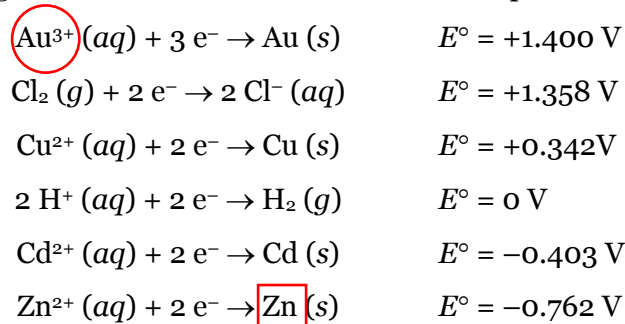


1. Consider the following half reactions and standard reduction potentials.



A) Circle (○) the strongest oxidizing agent and box (□) the strongest reducing agent.

B) Write the balanced chemical equation and cell diagram for the Galvanic cell with the largest standard cell potential (E_{cell}°).

The largest standard cell potentials would arise from: $2 \text{Au}^{3+}(\text{aq}) + 3 \text{Zn}(\text{s}) \rightarrow 2 \text{Au}(\text{s}) + 3 \text{Zn}^{2+}(\text{aq})$

$$E_{\text{cell}}^{\circ} = E^{\circ}(\text{Au}^{3+}/\text{Au}) - E^{\circ}(\text{Zn}^{2+}/\text{Zn}) = 1.400 \text{ V} - (-0.762 \text{ V}) = +2.162 \text{ V}$$

The cell (line) diagram would be: $\text{Zn}(\text{s}) \mid \text{Zn}^{2+}(1.00 \text{ M}) \parallel \text{Au}^{3+}(1.00 \text{ M}) \mid \text{Au}(\text{s})$

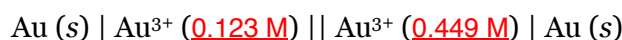
2. How long will it take to plate 0.0625 g of solid copper from an aqueous solution of copper(II) sulfate with a current of 0.200 A? Note: $1 \text{ A} = 1 \text{ C/s}$ and $F = 96,500 \text{ C/mol e}^{-}$

$$0.0625 \text{ g Cu} \times \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} \times \frac{2 \text{ mol e}^{-}}{1 \text{ mol Cu}} \times \frac{96500 \text{ C}}{1 \text{ mol e}^{-}} \times \frac{1 \text{ s}}{0.200 \text{ C}} = 949 \text{ s}$$

-or-

$$0.200 \frac{\text{C}}{\text{s}} \times \frac{1 \text{ mol e}^{-}}{96500 \text{ C}} \times \frac{1 \text{ mol Cu}}{2 \text{ mol e}^{-}} \times \frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} \times \frac{1}{0.0625 \text{ g Cu}} = 0.00105_4 \text{ s}^{-1} \rightarrow 949 \text{ s}$$

3. A Galvanic cell is constructed using two Au electrodes and two Au^{3+} solutions: one is 0.123 M and the other is 0.449 M. Fill in the concentrations in the cell diagram below for this cell?



4. You construct the following Galvanic/voltaic cell at 298.15 K.



What will the potential of the cell be after 0.10 M of Cr^{2+} is consumed?

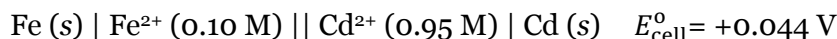
Assume that volume and temperature do not change.

	2 $\text{Cr}^{2+}(\text{aq})$	+	$\text{Co}^{2+}(\text{aq})$	\rightarrow	2 $\text{Cr}^{3+}(\text{aq})$	+	$\text{Co}(\text{s})$
I	0.30 M		0.20 M		2.00 M		n/a
C	-0.10		$-\frac{1}{2} \times 0.10$		+0.10		n/a
F	0.20		0.15		2.10		n/a

Now use the Nernst equation to find the new cell potential:

$$\begin{aligned} E_{\text{cell}} &= E_{\text{cell}}^{\circ} - \frac{RT}{nF} \ln \left(\frac{[\text{Cr}^{3+}]^2}{[\text{Cr}^{2+}]^2 [\text{Co}^{2+}]} \right) \\ &= 0.220 \text{ V} - \frac{\left(\frac{8.314 \text{ J}}{\text{mol} \cdot \text{K}} \right) (298.15 \text{ K})}{(2 \text{ mol e}^{-}) \left(\frac{96500 \text{ C}}{\text{mol e}^{-}} \right)} \cdot \ln \left(\frac{[2.10]^2}{[0.20]^2 \times [0.15]} \right) \\ E_{\text{cell}} &= 0.135 \text{ V} \end{aligned}$$

5. You construct the following Galvanic cell at 298.15 K.



The initial mass of the Fe electrode is 100.0 g and the volumes of the solutions are 1.00 L each. What will the cell potential be when the mass of the Fe electrode is 62.0 g? {Fe = 55.85 g/mol}

Assume that temperature does not change.

The net ionic equation is: $\text{Fe (s)} + \text{Cd}^{2+} (\text{aq}) \rightarrow \text{Fe}^{2+} (\text{aq}) + \text{Cd (s)}$

$$n_{\text{Fe}} = (100.0 - 62.0) \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} = 0.680_4 \text{ mol Fe consumed}$$

$$n_{\text{Cd}^{2+}} = 0.95 \text{ mol} - \left(0.680_4 \text{ mol Fe} \times \frac{1 \text{ mol Cd}^{2+}}{1 \text{ mol Fe}} \right) = 0.269_6 \text{ mol Cd}^{2+} \text{ leftover}$$

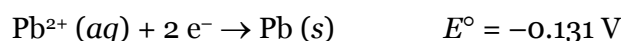
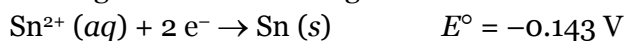
$$n_{\text{Fe}^{2+}} = 0.10 \text{ mol} + \left(0.680_4 \text{ mol Fe} \times \frac{1 \text{ mol Fe}^{2+}}{1 \text{ mol Fe}} \right) = 0.780_4 \text{ M Fe}^{2+} \text{ total}$$

Now use the Nernst equation to find the new cell potential:

$$E_{\text{cell}} = 0.044 \text{ V} - \frac{\left(8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}} \right) (298.15 \text{ K})}{(2 \text{ mol } e^{-}) \left(96500 \frac{\text{C}}{\text{mol } e^{-}} \right)} \cdot \ln \left(\frac{0.780_4 \text{ mol Fe}^{2+} / 1.00 \text{ L}}{0.269_6 \text{ mol Cd}^{2+} / 1.00 \text{ L}} \right)$$

$$E_{\text{cell}} = 0.030 \text{ V}$$

6. You construct a voltaic cell using the two reactions given.



If the cell starts with $[\text{Sn}^{2+}] = 1.35 \text{ M}$ and $[\text{Pb}^{2+}] = 2.11 \text{ M}$ at 298.15 K, what will be the concentration of Pb^{2+} when the cell is “dead”?

A “dead” cell is one that has reached equilibrium: $\Delta G = 0$ and $E_{\text{cell}} = 0 \text{ V}$.

	Sn (s)	+	Pb ²⁺ (aq)	→	Sn ²⁺ (aq)	+	Pb (s)
I	n/a		2.11 M		1.35 M		n/a
C	n/a		-x		+x		n/a
F	n/a		2.11 - x		1.35 + x		n/a

$$0 \text{ V} = 0.012 \text{ V} - \frac{\left(8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}} \right) (281 \text{ K})}{(2 \text{ mol } e^{-}) \left(96500 \frac{\text{C}}{\text{mol } e^{-}} \right)} \cdot \ln \left(\frac{1.35 + x}{2.11 - x} \right)$$

$$x = 1.17_4 \text{ M}$$

Now solve for the concentration of Pb^{2+}

$$[\text{Pb}^{2+}] = 2.11 - 1.17_4 = 0.94 \text{ M}$$

7. Consider a voltaic cell with the following cell diagram at 298.15 K.



- A) What will happen to the cell potential if $[\text{Pb}^{2+}]$ is doubled?

Increases

Decreases

Stays the same

- B) What will happen to the cell potential if $[\text{Cu}^{+}]$ is doubled?

Increases

Decreases

Stays the same

- C) What will happen to the cell potential if we added enough water to double the volumes of both the anodic and cathodic solutions?

Increases

Decreases

Stays the same